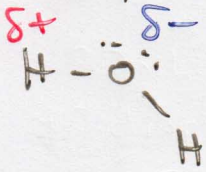
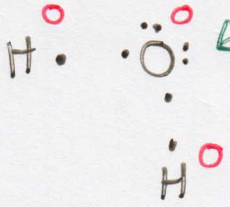
5/21/19

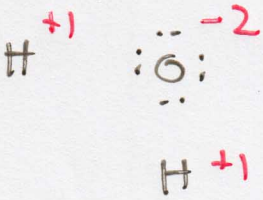


Formal Charge – Used in Lewis dot structures
→ bonds are assumed to be covalent

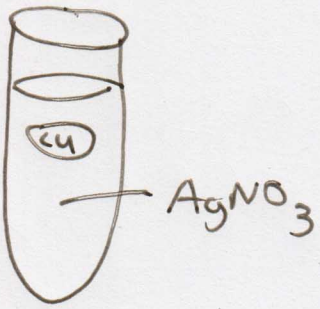


← When counting electrons, bonds are broken evenly (one electron goes to each atom in the bond),

Oxidation State – Used in redox reactions
→ all bonds are treated as ionic

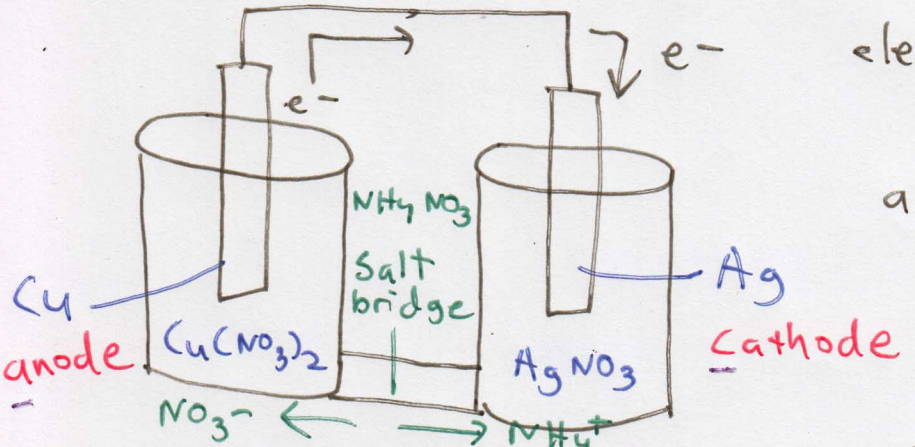


← All electrons in a bond are given to the more electronegative atom.



$Cu^{0}_{(s)} \longrightarrow Cu^{+2}_{(aq)} + 2e^{-}$

$2Ag^{+}_{(aq)} + 2e^{-} \longrightarrow 2Ag^{0}$



$Cu^{0} \rightarrow Cu^{+2}_{(aq)} + 2e^{-}$ $Ag^{+}_{(aq)} + e^{-} \rightarrow Ag^{0}$

oxidation reduction

electrode – a connection to an electrochemical cell

anode – the electrode at which oxidation occurs

cathode – the electrode at which reduction occurs

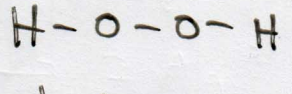
Salt bridge – contains a non-reactive salt, the ions of which are used to balance out the charges generated in the different parts of the cell.

spectator ion – an ion that is present but that does not participate in a reaction.

counter ion - a spectator ion used just to balance charge

Determining oxidation states O.S.

- any element in its unreacted form has an O.S. of zero (it hasn't yet gained or lost any electrons)
- the charge of a monatomic ion is the same as its oxidation state
- the sum of oxidation states is equal to the overall charge
- hydrogen normally has a +1 O.S.;
  oxygen normally has a -2 O.S.
  (except in peroxides where the O.S. is -1)



hydrogen peroxide
(O_2^{-2})

$\overset{+1}{Na}_2\overset{+6}{Cr}_2\overset{-2}{O}_7$

$2Cr + 7O = -2$

$2Cr + 7(-2) = -2$

$2Cr + -14 = -2$

$2Cr = +12$

$Cr = +6$